South Pasadena • AP Chemistry

Name______________________________
Period __ Date __/__/__

6 • Energy and Chemical Reactions

CALORIMETRY

Measuring heat (formerly measured in calories) is called **calorimetry**. Now we measure heat energy in Joules (J). The equation we use is:

$$q = m \cdot C \cdot \Delta T$$

q = heat energy
m = mass of water
C = the specific heat capacity
ΔT = the change in temperature (in °C or K)

1. Water has a specific heat capacity of 4.184 J/g·°C.
This means it takes 4.184 J to heat 1.00 gram of water 1.00°C.
   a) How much energy will it take to heat 10.0 grams of water 1°C? 41.84 J
   b) How much energy is needed to heat 30.0 g H_2O from 10.0 °C to 50.0 °C? 5020.8 J or 5020 J

$q = (30.0g)(4.184 \frac{J}{g\,°C})(40.0°C) =$

2. Let's try a standard **calorimetry** problem.
A pot of water (2.5 Liters of water) initially at 25.0°C is heated to boiling (100.°C). How much energy (in J) is needed to heat the water? (The density of water is 1 g/mL.)

$2.5L \times \frac{1000mL}{1L} \times \frac{1g}{1mL} = 2500g$

$q = (2500g)(4.184 \frac{J}{g\,°C})(75.0°C) = \boxed{784500 J = 780000 J}$

What would this amount of heat be in kJ? 784.5 kJ = $\boxed{780 kJ}$ 2 sig. figs.

3. What amount of heat is *released* when 175 g of water *cools* from 100.°C to room temperature, 20.0 °C?

$q = (175g)(4.184 \frac{J}{g\,°C})(-80.°C) = -58576 J = -58.6 kJ = \boxed{-59 kJ}$

⊖ sign mean heat is given off.

4. We don't always have to warm up or cool down water. The specific heat capacity of copper metal is 0.39 J/g·°C. It is easier (easier/~~more difficult~~) to heat up copper than to heat up water.

How much energy would it take to heat up a 5.20 g sample of copper from 20.0 °C to 100.°C?

$q = (5.20g)(.39 \frac{J}{g\,°C})(80.°C) = 162.24 J = \boxed{160 J}$

5. If 300. J of heat energy were used to heat up a 5.00 gram sample of copper metal and a 5.00 gram sample of water both starting at 10.0°C, calculate the final temperature of each sample?

Cu

$q = m\ c\ \Delta T$

$300J = (5.00g)(.39 \frac{J}{g\,°C})\,x$

$\Delta T = x = \frac{(300)}{(5)(.39)} = 154°$

$\boxed{T_f = 164°C}$

H_2O

$q = m\ C\ \Delta T$

$300J = (5.00g)(4.184 \frac{J}{g\,°C})\,x$

$\Delta T = x = \frac{(300)}{(5.00)(4.184)} = 14.3°$

$\boxed{T_f = 24.3°}$

Signs of ΔT and q:

– q means heat is **released**. + q means heat is **absorbed**.

ΔT is always ***final*** temperature – ***initial*** temperature.

If something is getting **hotter** (10° → 30°) the ΔT is 30 – 10 = + 20°. (heat is **absorbed**)

If something is getting **cooler** (75° → 25°) the ΔT is 25 – 75 = – 50°. (heat is **released**)

6. Suppose we mix 90.0 grams of **hot water** (90.0°C) with 10.0 grams of **cold water** (10.0°C).
Let x = the final temperature. C = 4.184 J/g·°C

a. Set up an expression for the energy ***released*** (q) by the hot water ($\Delta q_{hot} = m_{hot}C\Delta T_{hot}$)

$q = (90.0g)(4.184 \frac{J}{g°C})(x - 90.0°C)$

b. Set up an expression for the energy ***absorbed*** (q) by the cold water ($\Delta q_{cold} = m_{cold}C\Delta T_{cold}$)

$q = (10.0g)(4.184)(x - 10.0°C)$

c. Knowing that the heat released = – heat absorbed, combine the two expressions and solve for x.

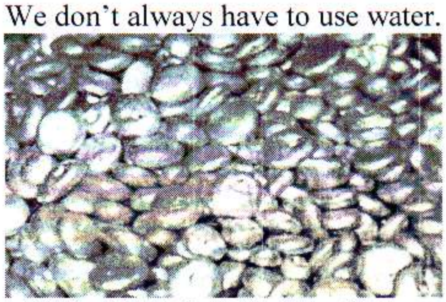
Hot $q = -q$ cold

$(90)(\cancel{4.184})(x - 90) = -(10)(\cancel{4.184})(x - 10)$

$90x - 8100 = -10x + 100$

$\boxed{x = 82\ °C}$

7. We don't always have to use water. Let's use some **aluminum shot**.



"shot" are these little pellets.

175 grams of hot aluminum (100.°C) is dropped into an insulated cup that contains 40.0 mL of ice cold water (0.0°C). Follow the example above to determine the final temperature, x.

a. Set up an expression for the heat lost by the aluminum (C=0.900 J/g·°C) $q = (175g)(.900 \frac{J}{g°C})(x - 100.°C)$

b. Set up an expression for the heat gained by the cold water.

$q = (40.0g)(4.184 \frac{J}{g°C})(x - 0.0°C)$

c. Put the two expressions together (don't forget to change one of the signs) and solve for x.

Al $q = -q_{H_2O}$

$(175)(.9)(x - 100) = -(40)(4.184)(x - 0)$

$157.5x - 15750 = -167.4x - 0$

$+167.4x \quad +15750 \qquad +167.4x \quad +15750$

$324.9x = 15750$

$\boxed{x = 48.5°}$

8. **Somewhat Confusing Definitions:**
There are several terms used in this chapter that sound very similar. Use the data provided to calculate each of them to clarify the differences. I've added some "Notes" that I hope will help.

74.8 J of heat is required to raise the temperature of 18.69 g of silver from 10.0°C to 27.0°C.

a. What is the **heat capacity** of the silver sample? (J/°C)
Note: This is a useful value only for this specific sample of silver.

$\Delta T = 27.0 - 10.0 = 17.0°C$

$\frac{74.8 J}{17.0°C} = \boxed{4.4 \frac{J}{°C}}$

b. What is the **specific heat capacity** of silver? (J/g·°C)
Note: This is a useful value for ***any*** sample of silver that is heated or cooled. This is equivalent to the 4.184 $J \cdot g^{-1} \cdot °C^{-1}$ that we use for water. This value is also called the **specific heat**.

$\frac{74.8 J}{(18.69g)(17.0°)} = \boxed{0.235 \frac{J}{g°C}}$